

10.7: Acid-Base Gallery

Learning Objectives

Make sure you thoroughly understand the following essential ideas which have been presented above. It is especially important that you know the precise meanings of all the highlighted terms in the context of this topic.

- Describe some of the special properties of *sulfuric acid* that make it especially important both in the laboratory and in industry.
- Name the major acids and bases that are important to the **fertilizer** industry
- Name the natural sources of any three of the major **organic** acids.
- What are *fatty acids*? In what major way do the physical properties of saturated and unsaturated fatty acids differ?
- Describe the general structure of an *amino acid*, and state why they are important.

Acids and bases are of interest not only to the chemically inclined; they play a major role in our modern industrial society — so anyone who participates in it, or who is interested in its history and development, needs to know something about them. Five of the major acids and bases fall into the "Top 20" industrial chemicals manufactured in the world. The following table shows year-2000 figures for the U.S:

cal and rank	Sulfuric acid - 1	Lime (CaO) 3	Phosphoric acid - 4	Ammonia 5	Sodium hydroxide - 9	Nitric acid - 11
production in 10 ⁹ kg	40	20	16	15	11	8
major use	chemicals	cement	fertilizers	fertilizers	chemicals	chemicals

The mineral acids

This term refers to any inorganic acid, but its common use is usually limited to the major strong acids plus phosphoric acid. The major mineral acids— sulfuric, nitric, and hydrochloric— have been known since medieval times. Their discovery is usually credited to the Persian alchemist Abu Musa Jabir ibn Hayyan, known in the West by his Latinized name Geber. Jabir also invented *aqua regia*, the mixture of nitric and hydrochloric acids that has the unique ability to dissolve gold.

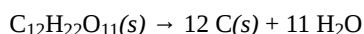
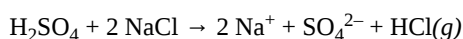
Sulfuric acid

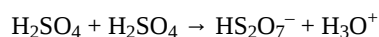
More sulfuric acid is manufactured than any other industrial chemical, and it is the cheapest industrial acid worldwide. It has been continuously manufactured in the U.S. since 1793 and in Europe for much longer.



What you should know about it

- Pure anhydrous H₂SO₄ is a dense, viscous liquid which melts at 10.4°C and boils at about 300°C, decomposing back into its constituents, H₂O and SO₃.
- The acid undergoes **autoprotolysis**





- Its **high boiling point** makes the acid ideal for making other acids, such as nitric and hydrochloric, which are more volatile; removal of the gaseous product drives the reaction to the right as predicted by the *Le Chatelier Principle*:
- Sulfuric acid has a voracious appetite for water, and thus is an excellent **dehydrating agent**. This is seen most spectacularly if some concentrated acid is poured onto a small pile of table sugar; after a short time, a vigorous reaction ensues resulting in a pile of porous, steaming carbon
- Sulfuric acid can even dehydrate itself!
, some of which have been detected on the surface of Jupiter's moon Europa.
- Owing to the autoprotolysis and self-dehydration reactions described above, "pure" sulfuric acid contains at least six minority species in addition to H_2SO_4 .



How it is made

Sulfur trioxide, the anhydride of sulfuric acid, is the immediate precursor. Gaseous SO_3 reacts vigorously with water, liberating much heat in the process:



Industrial manufacture of the acid starts with sulfur dioxide, prepared from burning elemental sulfur or obtained as a byproduct from roasting sulfide ores. The oxidation of SO_2 to SO_3 looks simple



but there are several complications:

- All chemical reactions take place more rapidly at higher temperatures, but because this reaction is highly exothermic, raising the temperature decreases the yield.
- Because of the decrease in volume (1.5 moles of gases to 1 mole), raising the pressure will increase the yield, so the reaction is carried out at a temperature below 600°C but at very high pressure.
- Specialized catalysts are used to speed up the reaction at these lower temperatures.
- Dissolving the SO_3 directly in water would release large amounts of heat, creating a mist of fine acid droplets that would escape into the atmosphere. The SO_3 is instead dissolved in sulfuric acid to form *pyrosulfuric acid* or *oleum*, sometimes known as *fuming sulfuric acid*:



- The oleum is then treated with water to form industrial grade (96-98%) sulfuric acid:

What it is used for

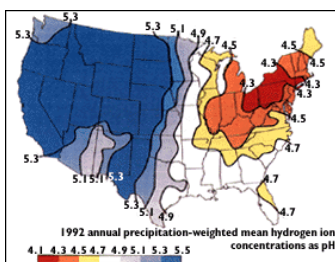
Sulfuric acid has a broad spectrum of industrial uses, and the annual tonnage follows the economic cycle quite closely.

- Sixty percent of worldwide product goes into the manufacture of *phosphoric acid* H_3PO_4 which is used to make phosphate fertilizers and phosphate-based household detergents.
- An important nitrogen fertilizer, ammonium sulfate $(\text{NH}_4)_2\text{SO}_4$, is made by reacting sulfuric acid with ammonia, the latter often obtained from the thermal decomposition of coal.

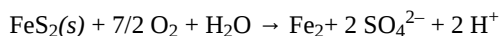
- Sulfuric acid is the major component of pickling acid that is used to remove surface oxide scale from steel before it is fabricated into steel product for the automobile and other industries.
- Aluminum sulfate (made from bauxite Al_2O_3 with H_2SO_4), is widely used in the papermaking industry to coagulate the cellulose fibers, producing a smooth, hard paper surface. Another major use is to make aluminum hydroxide which is used to filter out particulate matter in water treatment facilities.
- The familiar lead-acid storage battery employs sulfuric acid as an electrolyte. As the battery discharges, the concentration of sulfuric acid in the electrolyte decreases as sulfate ions are taken up as PbSO_4 . Owing to the high density of the acid, the state of charge of the battery can be measured by means of a hydrometer.

Sulfuric acid in the environment

Acid Rain - Combustion of fossil fuels which contain organic sulfur compounds releases SO_2 into the atmosphere. Photochemical oxidation of this compound to SO_3 , which rapidly takes up moisture, leads to the formation of H_2SO_4 , a major component of acid rain.



Acid mine drainage results when sediments of the very common *iron pyrite*, FeS_2 , are exposed to air and are oxidized:



further oxidation of the iron to Fe^{3+} results in additional reactions. The resulting drainage liquid is often orange-brown in color and can have a pH of below zero.

Nitric acid

Anhydrous HNO_3 is a colorless liquid boiling at 82.6°C , but "pure" HNO_3 only exists as the solid which melts at -41.6°C . In its liquid and gaseous states, the acid is always partially decomposed into nitrogen dioxide:



This reaction, which is catalyzed by light, accounts for the brownish color of HNO_3 solutions.

What you should know about it

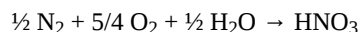
- HNO_2 undergoes autoprotolysis to a greater extent than any other liquid. Further reactions of the conjugate acid H_2NO_3^+ with HNO_3 lead to a complicated mixture of species in the liquid.
- Dilute nitric acid can be concentrated by distillation up to a maximum of 68%, at which point it forms a constant-boiling (azeotropic) mixture with water. Higher concentrations require dehydration with sulfuric acid; the result is fuming nitric acid.
- "Concentrated nitric acid" is sold as a 70% solution in water, corresponding to a concentration of about 16M.
- Nitric acid is a very strong **oxidizing agent**, which adds to its corrosive behavior with organic materials including, of course, skin, which it turns yellow owing to a reaction with the protein *keratin*. Reactions with many organic compounds are highly exothermic and often violent. The well-known reaction of nitric acid with metallic copper produces copious amounts of brown nitrogen dioxide gas.



<http://www.angelo.edu/> [image link]

How it is made

The simplest method, which was used industrially before 1900, was by treatment of sodium nitrate ("Chile saltpeter", NaNO_3) with sulfuric acid. The direct synthesis of the acid from atmospheric nitrogen and oxygen is thermodynamically favorable



but is kinetically hindered by an extremely high activation energy, a fact for which we can be most thankful (see sidebar.) The first industrial nitrogen fixation process, developed in 1903, used this reaction to produce nitric acid, but it required the use of an electric arc to supply the activation energy and was therefore too energy-intensive to be economical.

If it were not for the high activation energy required to sustain this reaction, all of the oxygen in the atmosphere would be consumed and the oceans would be a dilute solution of nitric acid.

The modern Ostwald process involves the catalytic oxidation of ammonia to nitric oxide NO , which is oxidized in a further step to NO_2 ; reaction of the latter with water yields HNO_3 . This route, first developed in 1901, did not become practical until the large-scale production of ammonia by the [Haber-Bosch process](#) in 1910.

What it is used for

The major industrial uses of nitric acid are for the production of ammonium nitrate **fertilizer**, and in the manufacture of **explosives**. On a much small scale, the acid is used in metal pickling, etching semiconductors and electronic circuit boards, and in the manufacture of acrylic fibers.

In the laboratory, the acid finds use in a wide variety of roles.

In the environment

High-temperature combustion processes (in internal combustion engines, power plants, and incinerators) can oxidize atmospheric nitrogen to nitric oxide (NO) and other oxides (" NO_x "); the NO is then photooxidized to NO_2 , which reacts with water to form HNO_3 which is a major component of [acid rain](#). NO_2 is the major precursor of [photochemical smog](#).

Hydrochloric acid

Unlike the other major acids, there is no such substance as "pure" hydrochloric acid; what we call "hydrochloric acid" is just an aqueous solution of hydrogen chloride gas (bp -84°C). But in a sense it is more "pure" than the acids discussed above, since there is no autoprotolysis; hydronium and chloride ions are the only significant species in the solution. Hydrochloric acid is usually sold as a 32-38% (12M) solution of HCl in water; concentrations greater than this are known as *fuming hydrochloric acid*.

Hydrochloric acid is still sometimes sold under its older name muriatic acid for cleaning bricks and other household purposes. The name comes from the same root as marine, reflecting its preparation from salt.

The acid has been known to chemists (and alchemists), and used for industrial purposes since the middle ages. Its composition HCl was demonstrated by Humphrey Davy in 1816.

What you should know about it

- Hydrochloric acid is the least hazardous of the strong mineral acids to work with because unlike the other ones, it is not an oxidizing agent. It is usually the acid of choice for titrations and other operations in which the main requirement is simply a strong source of hydronium ions.
- The concentrated acid boils at 48°C . As boiling continues, it loses HCl and the boiling point rises to 109°C , at which point a constant-boiling (azeotropic) solution remains, consisting of 20.2% HCl .



What it is used for

The uses of hydrochloric acid are far too many to enumerate individually, but the following stand out:

- A major industrial use is to **remove surface scale** from iron or steel ("pickling") before it is processed into sheets or other forms, or galvanized or coated.
- Production of **chlorinated organic chemicals**, particularly vinyl chloride, polyurethanes, and other construction polymers, consumes huge amounts of HCl.
- The acid is widely used for pH control of water, including neutralization of wastewater streams, and for regenerating **ion-exchange water softeners**.

How it is made

The ancient method of treating salt with sulfuric acid to release HCl has long since been supplanted by more efficient processes, including direct synthesis by "burning" hydrogen gas in chlorine:



Most hydrochloric acid production now comes from reclaiming byproduct hydrogen chloride gas from other processes, especially those associated with the production of industrial organic compounds.

2 The Alkali Metals

The term *alkali* usually means a basic salt of a **Group 1 or 2** ("alkali" or "alkaline earth") metal. All alkalis are of course bases, but the latter term is much more general, whether defined according to the Arrhenius, Brønsted-Lowry, or Lewis concepts. The word *alkali* comes from the Arabic *al-qali*, which refers to the ashes from which sodium and potassium hydroxides (*potash*, "ashes remaining in the pot", and the origin of the element name *potassium*) were extracted as a step in the making of soap.

Sodium hydroxide

Pure sodium hydroxide is a white solid consisting of Na^+ and OH^- ions in a crystal lattice. Although it is widely thought of as an ionic solid, *van der Waals forces* make a substantial contribution to its stability.

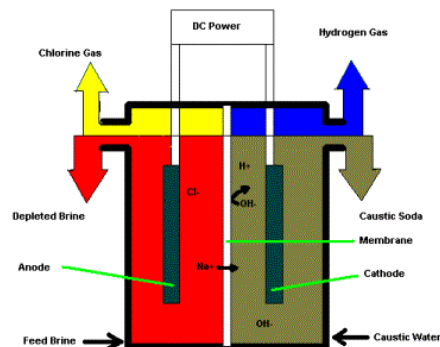


What you should know about it

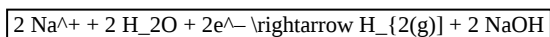
- Owing to its *deliquescence* (ability to take up moisture) and its tendency to react with carbon dioxide, the solid must be stored in a closed container.
- In industry, sodium hydroxide is commonly known as *caustic soda* or simply as **caustic**; NaOH sold for household purposes is usually known as *lye*.
- Sodium hydroxide slowly attacks **glass** to form sodium silicate. Glass vessels used to store concentrated solutions gradually develop a cloudy coating on the inside.
- Some **metals**, notably aluminum, zinc, and titanium, react with strongly alkaline solutions, but iron and copper are immune to this kind of attack.
- Highly alkaline solutions also soften and dissolve **skin**, accounting for the slippery feeling associated with strong bases. Sodium hydroxide was once used to dispose of animal carcasses, digesting them into an easily disposable liquid form.

How it is made

Sodium hydroxide is now manufactured by the electrolysis of brine solutions, and along with chlorine, is one of the two major products of the *chloralkali industry*.



Electrolysis of aqueous NaCl produces Cl₂ at the anode, but because H₂O can be reduced more readily than Na⁺, the water is decomposed to H₂ and OH⁻ at the cathode, leaving a solution of NaOH. An older *mercury cell process* reduces the Na⁺ to Na within a mercury amalgam (alloy), and the metallic sodium is then combined with water to produce NaOH and hydrogen. The net reaction for the reduction step is the same for both methods:



The resulting solution is usually evaporated to such a high concentration that it solidifies at ordinary temperatures. It is commonly shipped in rail cars or barges which can be heated with steam to liquefy the mixture for removal. (It is obviously uneconomical to ship large quantities of water across the country!)

What it is used for

- Sodium hydroxide is one of the most diverse industrial chemicals in terms of its applications. Most householders know it as the active ingredient of **drain cleaning agents**.
- Huge quantities are consumed by the **pulp and paper** industry, which is probably its single largest specific industrial application. It is used to remove the lignin component of wood pulp from the cellulose so that the latter can be processed into paper.
- About half of the NaOH output goes into the production of a wide variety of other **industrial chemicals**, and in degreasing steel drums and other industrial surfaces.
- "Lye" plays a role in the processing of many types of **foods**, including chocolate, olives, pretzels, and the "hominy" and "grits" corn products used in the Southern U.S. It also acts as a chemical peeling agent for fruits and vegetables.



In contrast to the extremely diverse applications of sodium hydroxide which makes the demand for this commodity relatively immune to the ups and downs of the economic cycle, the consumption of chlorine is directly dependent on the economy as reflected in the demand for polyvinyl chloride products that are now widely used in the construction and home furnishings industries. Because chlorine, being a gas, is expensive to store, the output of the chloralkali industry is governed largely by the demand for this commodity. When times are good this presents no problem; caustic is then largely a by-product and can easily be stockpiled if supply exceeds demand. But during an economic downturn, the demand for chlorine declines, limiting its production along with that of caustic. But because the demand for caustic tends to decline much less, it becomes scarce and its price rises, thus tending to drive the industrial economy into even deeper trouble.

Sodium carbonate

This compound is known industrially as **soda ash**, and domestically as *washing soda*. The common form is the *heptahydrate*, $\text{Na}_2\text{CO}_3 \cdot 7 \text{H}_2\text{O}$. The white crystals of this substance spontaneously lose water (*effloresce*) when exposed to the air, forming the monohydrate.

What it is used for

- Although carbonates are much weaker bases than hydroxides, a solution of sodium carbonate can still have a pH of 11 or so, sufficiently high to allow it to substitute for sodium hydroxide in many applications—especially when the price of caustic is high.
- The single most important use of soda ash is in the manufacture of glass, where it serves to lower the melting point of the principal component, SiO_2 .
- Another emerging major use is to neutralize the SO_2 emissions of fossil fuel-burning power plants.
- The older use of sodium carbonate as a **cleaning agent** (hence the name *washing soda*) was based partly on the ability of its alkaline solutions to emulsify grease, but mainly as a means of precipitating the insoluble carbonates of calcium and magnesium before these ions (commonly present in hard water) could form undesirable precipitates with soaps. The use of modern detergents has largely eliminated this once important market.



How it is made

Most of the world's sodium carbonate is made by the ammonia-soda "Solvay" process developed in 1861 by the Belgian chemist Ernest Solvay (1838-1922) whose patents made him into a major industrialist and a rich philanthropist. This process involves a set of simple reactions that essentially converts limestone (CaCO_3), ammonia NH_3 and brine (NaCl) into sodium bicarbonate NaHCO_3 and eventually Na_2CO_3 , recycling several of the intermediate products in an ingenious way.



A minor source of soda ash (but quite significant in some countries, such as the U.S.) is the mining of natural *evaporites* (the remains of ancient lakes), such as the *trona* found in Southern California.

Ammonia

Ammonia NH_3 is of course not a true alkali, but it is conveniently included in this section for discussion purposes. Most people are familiar with the pungent odor of this gas, which can be detected at concentrations as low as 20-50 ppm.



Tradition dies slowly: a non-existent chemical available in bottles!

What you should know about it

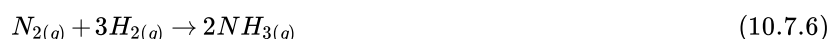
- More moles of ammonia are manufactured than of any other industrial chemical.
- Ammonia is extremely soluble in water. An aqueous solution of ammonia is still sometimes referred to in commerce as "ammonium hydroxide", but this term is no longer favored by chemists because no such compound as NH_4OH has ever been shown to exist. At neutral pH, about 99% of the ammonia in water exists as NH_4^+ ions.
- Ammonia is an end product of nitrogen metabolism in most organisms. One source that may be familiar to parents of infants is the bacterial decomposition of the contents of diapers.
- Liquid ammonia (bp -33°C) is often used as an ionizing laboratory solvent.

What it is used for

- The major use of ammonia (about 80%) is as a fertilizer, most commonly as anhydrous ammonia (the gas is injected directly into the soil) or after conversion to $(\text{NH}_4)_2\text{SO}_4$, NH_4NO_3 , or urea $\text{O}-\text{C}(\text{NH}_3)_2$.
- Ammonia is used in the production of numerous polymers, including nylons and polyurethanes.
- Explosives manufacture accounts for about 5% of ammonia production.
- Beyond these, there are hundreds of minor uses, including as a household cleaning agent (aqua ammonia), a refrigerant, and as a laboratory reagent.

How it is made

Ammonia is made by direct synthesis from the elements:



... a simple-looking reaction, but one that required some very creative work to implement; the *Haber-Bosch process* is considered to be the most important chemical synthesis of the 20th Century.

3 Some important organic acids

Most acids are organic— there are millions of them. The acidic function is usually a hydroxyl group connected to a carbon that is bonded to an electron-withdrawing oxygen atom; the combination is the well-known *carboxyl group*, $-\text{COOH}$. Here are a few that are part of everyone's life.

Acetic acid

This is next to formic acid in being the simplest of the organic acids, and in the form of vinegar (a 5-8% solution in water) its characteristic odor is known to everyone. The pure acid is a colorless liquid above 16.7°C ; below this temperature it forms a crystalline solid, hence the term "glacial acetic acid" that is commonly applied to the pure substance. The name of the acid comes from *acetum*, the Latin word for vinegar.



What you should know about it

- The pure acid, although quite weak in the proton-donor sense, is quite corrosive and its vapors are very irritating.
- A 1.0M solution of the acid has a pH of about 2.4, corresponding to only four out of every thousand CH_3COOH molecules being dissociated.

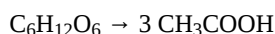
What it is used for

Slightly less than half of the world production of acetic acid goes into the production of **polymers**. The end product visible to most people would be the flexible plastic bottles in which drinking water is sold. Other uses are related mostly to the production of other chemicals, mainly *acetic anhydride*, but also including *aspirin*.

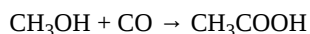


How it is made

Bacterial fermentation of sugars has been the source of vinegar since ancient times, and is still accounts for most food-grade acetic acid and vinegar, but it now amounts to only about 10% of total acetic acid production:



There are several important synthetic routes to acetic acid production, but the major one is by treating methanol with carbon monoxide:



A brief look at some other interesting organic acids



$\text{H}-\text{C}(=\text{O})-\text{OH}$ Formic acid
 (Methanoic acid) HCOOH
 mp 8.4°C.

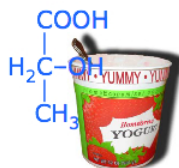
This, the simplest of the carboxylic acids, is the chemical weapon that Nature has given ants and bees (the Latin word for ant is *formicus*.) Known since the 15th Century, it was first distilled from ants, but is now made synthetically. Its main uses are as a preservative and antibacterial agent in livestock feed. Chemists find it a useful source of carbon monoxide (just add sulfuric acid.)



$\text{HOOC}-\text{COOH}$ Oxalic acid

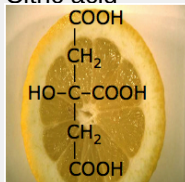
Just two carboxyl groups joined together, this acid has a special ability to grab up divalent metal ions and bind them into five-membered MO_2C_2 rings. Because many such metal ions (Ca^{2+} , for example) are essential nutrients, this can be dangerous to your health. The acid occurs in many plants, notably rhubarb (it is what makes the leaves poisonous), parsley and spinach. If you have ever noticed a funny feeling in your mouth after drinking milk with a rhubarb desert, it is due to precipitation of calcium oxalate. This same solid is often a major component of kidney stones, and it contributes to the miseries of gout.

Lactic acid



Structurally, lactic acid is both an alcohol and a carboxylic acid— not all that uncommon. We know it as the acid found in sour milk, yogurt, and in tired muscles. When the blood cannot deliver enough oxygen to oxidize glucose all the way to CO_2 and water, your muscles go into a much-less-energy-efficient "anaerobic" mode that generates lactate. In milk products, lactic acid is produced in about the same way by *acidophilus* bacteria that eat the *lactose* ("milk sugar") and who don't know how to utilize oxygen.

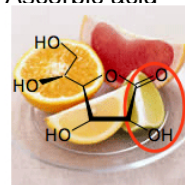
Citric acid



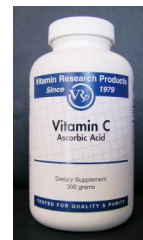
Known by Arabic alchemists in the 8th Century; first isolated by Scheele in 1784

Although it is one part alcohol and three parts acid, citric acid is quite weak, but nevertheless strong enough to make it unpleasant to suck on a lemon, of which it can comprise as much as 8% of the dry weight of this fruit. Biochemists know it as part of the *citric acid cycle*, a sequence of reactions involved in extracting energy from the oxidative metabolism of foods. Its major industrial use is as a food flavoring and preservative agent; large quantities are used to make soft drink beverages. The acid also finds use in cleaning agents.

Ascorbic acid

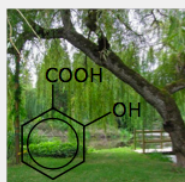


No carboxyl groups here, but the -OH group one carbon away from the double bond is still fairly acidic. One of its geometric isomers, L-ascorbic acid, is more widely known as **Vitamin C**; discovery of the essential role of this substance in preventing the disease *scurvy* (from which its name is derived) yielded two Nobel prizes in 1937. Most animals are capable of synthesizing their own Vitamin C, but primates (along with guinea pigs) seem to have lost the required gene somewhere along the way, so we must depend on fruits and veggies for our supply. (Because ascorbic acid is soluble in water, it tends to be leached out of vegetables when they are boiled, so it is much healthier to steam them.)



The major industrial use of ascorbic acid is as an *antioxidant*, so it is often added to foods and other materials as a preservative.

Salicylic acid



2-hydroxybenzoic acid is found in willow trees

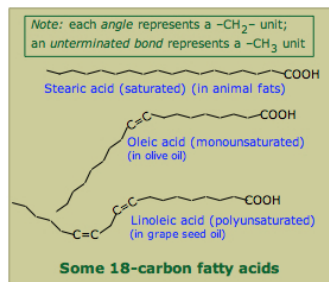
Does chemistry give you a headache? If so, pull a small shoot off of a willow tree and chew on it for a while. This has been a folk-remedy for pains and fever since ancient times. Willow bark (*salix* in Latin) contains, and lent its name to the active principle, salicylic acid.

The acid itself turned out to be a bit too much of an irritant to the stomach's lining, so the German firm Bayer began marketing a tamed version, *acetylsalicylic acid* (ASA) under the name Aspirin in 1899, and it has been going strong ever since. Interestingly, the detailed chemistry of its pain-and-fever reliving action was not discovered until 1970.



Fatty acids

This generic term refers to carboxylic acids built from a chain of 4 to 22 carbon atoms. Fatty acids, as the name implies, are derived from fats, in which they are bound to glycerol in the form of triglycerides. Fats, which occur in all animal tissues, are the most efficient means of storing metabolic energy. Vegetable oils are another source.



There are two general categories of fatty acids:

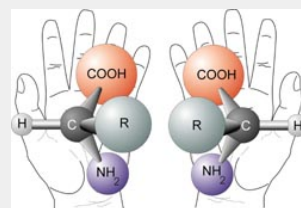
- **Saturated fatty acids** consist of straight chains of carbon atoms with a methyl group on one end and a carboxyl (acidic) group on the other, with *methylene groups* $\text{-CH}_2\text{-}$ in between. Thus stearic (octadecanoic) acid is $\text{CH}_3(\text{CH}_2)_{16}\text{COOH}$. The linear shape of these molecules allows them to pack together very efficiently, contributing to their ability to store energy in a small space; the higher saturated acids tend to be waxy solids.
- **Unsaturated fatty acids** contain one or more carbon-carbon double bonds, which introduces a complication: because free rotation around the axis of a double bond is not possible, the neighboring hydrogens can be either on the same side of the bond (*cis*) or on opposite sites (*trans*)—thus the *cis*- and *trans*-fatty acids. These double bonds introduce kinks into the carbon chain, especially in the case of the *cis*-acids. The bent and curved carbon chains that result cannot pack together compactly enough to interact very strongly with each other, so unsaturated fatty acids tend to be liquids.

The human body is able to synthesize most of the fatty acids it needs, but two classes of unsaturated acids, known as the **essential fatty acids**, can be obtained only from foods. The compounds are known as ω -3 and ω -6 fatty acids; the ω (omega)- n notation means that the double bond is located n carbons away from the terminal CH_3 group of the molecule.

Most of the unsaturated fatty acids found in nature have *cis* configurations around their double bonds. The *trans* fatty acids in our foods that we hear much about are made from vegetable oils that have had some of their double bonds changed to single bonds by a chemical process called hydrogenation. The object of this is to make the original liquid fatty acids into solid forms that are more suitable for use in foods (as in margarine) and particularly as shortenings in baking. Unfortunately, the hydrogenation process also changes the remaining *cis* double bonds into their *trans* forms which are believed to be implicated in cardiac disease.

Amino acids

Amino acids are the building blocks of proteins, which is what you are largely made of. There are twenty of them, of which about half can be manufactured by the body; we must depend on our food to supply the remainder, which are known as the *essential amino acids*. Each one comes in two mirror-image forms; only the "left" forms are found in most proteins.

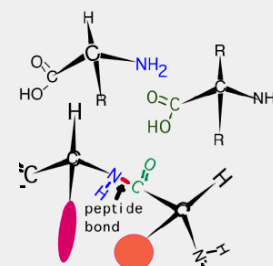


All amino acids have the basic structure shown above; they differ in the nature of the group of atoms designated by "R" in the diagram.

Although we call them "acids", the amino acids are really chemical hermaphrodites; you will recall that amines are weak bases. The balance between their acidic and basic properties can be shifted simply by changing the pH.

The carboxylic acid part of one amino acid can react with the amine part of another to form a peptide bond (an *amide* linkage) shown here.

The polymeric chains that result are known as peptides if they are fairly short (2 to about 20-50 amino acid units); longer chains, or aggregate made up of multiple peptide units, are proteins. Owing to their very large size (500 amino acid residues is quite common, but some have as many as 1500), proteins are able to fold in various ways, so the amino acid sequence alone is not sufficient to determine their properties.



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